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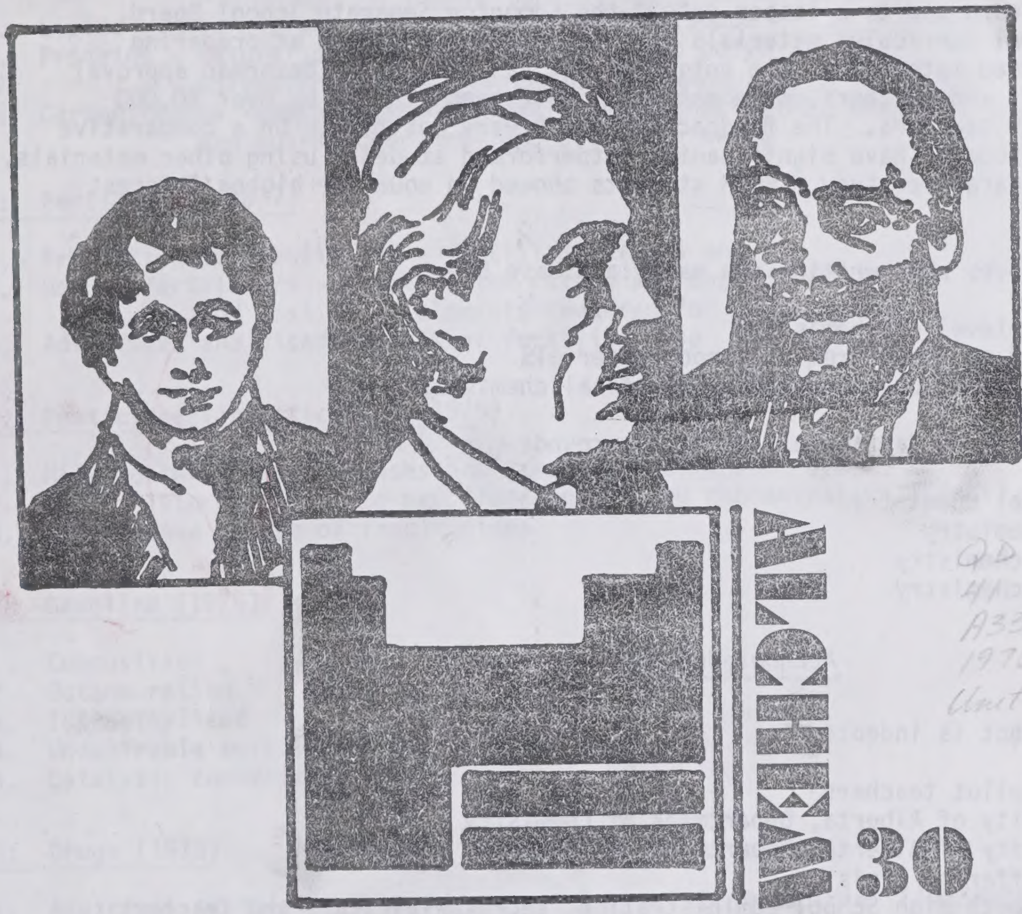

alberta chemistry project materials

FOREWORD

ALCHEM (Alberta Chemistry Project Materials) was initiated by the Edmonton Regional Chemistry Council. Funding has been provided by the Edmonton Public School Board and to a lesser extent the Edmonton Separate School Board. ALCHEM is a local curriculum materials development project aimed at preparing classroom oriented materials. The unique format has received widespread approval both by students and teachers. The materials have been piloted by over 20,000 students and 100 teachers. The feedback has been very positive. On a comparative study, ALCHEM students have significantly outperformed students using other materials. On the same comparative study ALCHEM students showed an equal or higher interest in chemistry.

The main objectives when writing the materials have been:

1. keep a high level of chemistry
2. provide easy-to-learn-from classroom materials
3. integrate applied chemistry and theoretical chemistry

The applied chemistry has largely centered around:

1. environmental chemistry
2. consumer chemistry
3. industrial chemistry
4. history of chemistry

Acknowledgements

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1. the ALCHEM pilot teachers
2. the University of Alberta, Department of Chemistry
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4. our much suffering typists
5. Queen Elizabeth High School administration, secretarial staff and teachers
6. the ALCHEMists - Frank Jenkins, George Klimiuk, Dean Hunt, Dick Tompkins, Oliver Lantz, Michael Dzwiniel, Dale Jackson, Tom Mowat, Myron Baziuk and Eugene Kuzub

First edition (1975)

Second (revised) edition (1976)

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ALCHEM 30 ELECTIVE UNITS

Elective Units

Unit O: Foods and Their Analogs (1975)

1. Lipids - natural fats and synthetic analogs (soaps and detergents)
2. Proteins - natural polypeptides and synthetic polymers (polyalkenes, polyesters and polyamides)
3. Carbohydrates - natural saccharides (monosaccharides, disaccharides and polysaccharides) and synthetic sweeteners

Unit P: Fertilizers (1977)

1. Production of fertilizers - fertilizer plants and the nitrogen cycle
2. Use of fertilizers - quantitative rating and application of fertilizers and listing of elements required for plant growth
3. Advantages and disadvantages of fertilizer use

Unit Q: Pesticides (Insecticides) (1975)

1. Main concerns about classes and structures of insecticides
2. DDT - history, threat to man (food chains and concentration effect)
3. Alternative to use of insecticides

Unit R: Gasoline (1975)

1. Composition
2. Octane rating
3. Tetraethyllead
4. Undesirable emissions
5. Catalytic converters

Unit S: Drugs (1975)

1. Aspirin - synthesis, % yield and melting point of aspirin (*Incomplete*)

Unit OS: Alberta's Oil Sands (1976)

1. Relative importance and extent of energy resources
2. Geography and geology
3. Composition of Athabasca deposit
4. Composition of bitumen
5. History of process research
6. Extraction of bitumen
7. Upgrading of bitumen
8. Ancillary reaction units
9. Future development

Future ALCHEM Elective Units (1977)

1. Alberta Chemical Industries (ALCHEM 10 or 20)
2. Sulfur
3. Petrochemicals
4. Petroleum
5. Coal

ALCHEM 30 - CORE UNITS

Unit K: Review of Chemistry 10 & 20

1. Elements and the periodic table
2. Compounds, bonding and nomenclature
3. Chemical reactions
4. The mole
5. Gravimetric stoichiometry
6. Chemical bonding (Lewis diagrams & VSEPR)
7. Organic chemistry
8. Solution chemistry
9. Solution stoichiometry

Unit L: Energy Changes

1. Force fields, bonds, energy and measuring energy changes
2. Phase changes - states of matter and molar heats of phase changes
3. Chemical changes - energy stored in bonds, endothermic and exothermic reactions, heats of formation, and molar heats of reaction
4. Nuclear changes - a comparison of the energy involved in phase, chemical and nuclear changes
5. Alternative energy sources

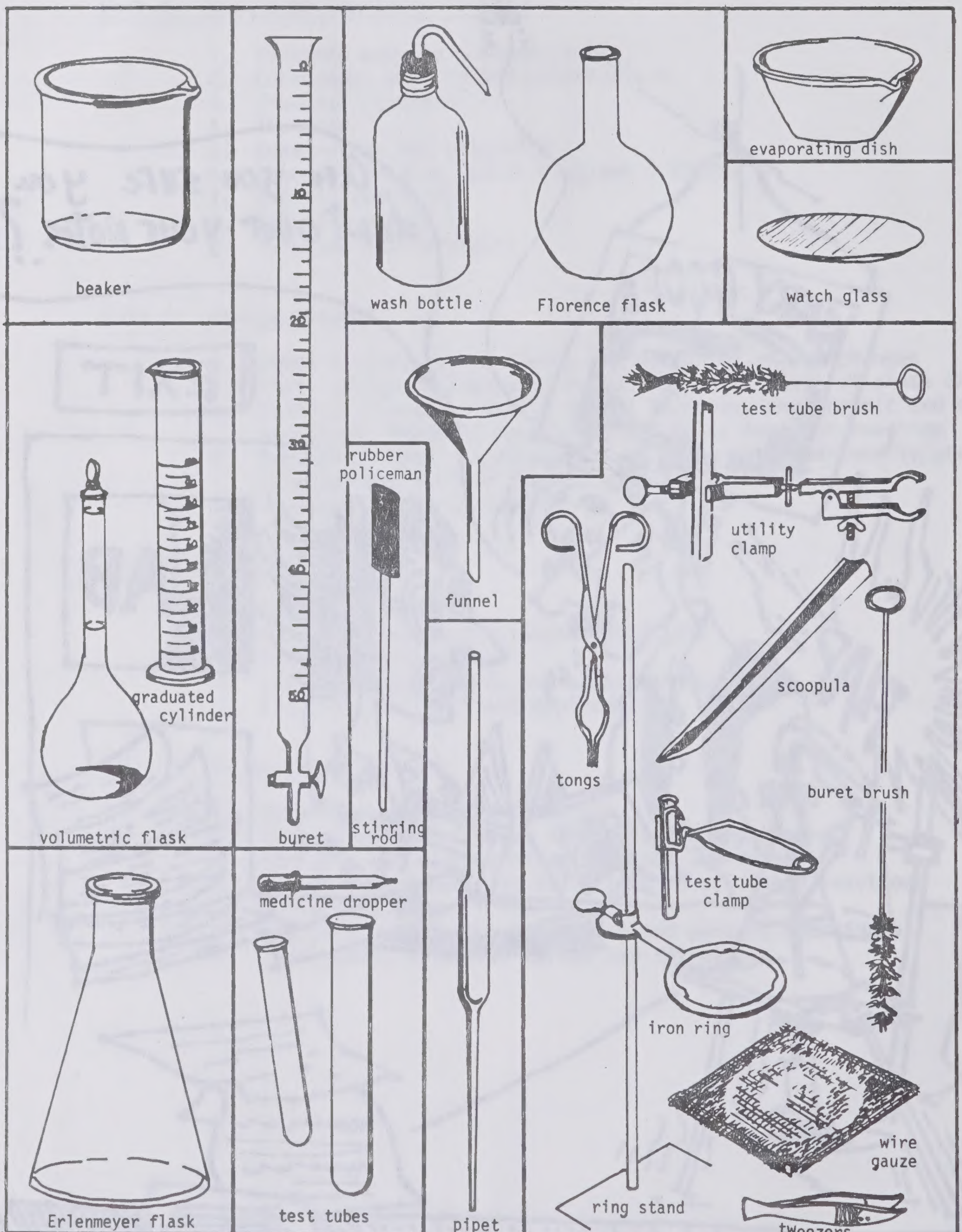
Unit M: Electrochemistry

1. Redox in a beaker
2. Half reactions and balancing equations
3. Reduction potentials
4. Electrochemical cells and applications
5. Electrolytic cells and applications

Unit N: Acids and Bases

1. Definitions and classification
2. Special nature of water in Bronsted-Lowry theory
3. Strength vs. concentration of acids and bases
4. Predicting products in and extent of acid-base reactions
5. Polyprotic acids
6. pH, $[H^+]$ and $[OH^-]$ calculations and percent reaction
7. Indicators, titration curves and titrations

UNIT K
Review of Chemistry 10 & 20


beaker
wash bottle
Florence flask
evaporating dish
watch glass
test tube brush
rubber policeman
funnel
utility clamp
graduated cylinder
volumetric flask
buret
stirring rod
tongs
scoopula
buret brush
medicine dropper
test tube clamp
iron ring
wire gauze
Erlenmeyer flask
test tubes
pipet
ring stand
tweezers

REVIEW OF CHEMISTRY 10 & 20

ATOMIC STRUCTURE AND CHEMICAL SPECIES

Definitions

nucleus - the central region of the atom

proton - a heavy, stable, positively charged particle found in the nucleus of all atoms

neutron - a heavy, stable, uncharged particle found in the nucleus of atoms

electron - a light, stable, negatively charged particle found in the region surrounding the nucleus

atom - an extremely stable arrangement of protons, neutrons and electrons. An atom is the smallest part of matter that can enter into chemical combination.

electron energy level - electrons around an atom possess specific amounts of energy and are most likely to be found in specific regions around the nucleus called energy levels. The number of electrons that can occupy any energy level is reflected in the periodic table of elements.

Bohr diagram of an atom - a diagram of an atom showing the number of protons and neutrons in the nucleus and the number of electrons in each energy level. Sometimes neutrons are not shown.

atomic number - the number of protons in the nucleus of an atom

mass number - the number of protons plus neutrons in the nucleus of an atom or simple ion

isotopes - two atoms that have the same number of protons but a different number of neutrons

valence electrons - the electrons in the outer most energy level of an atom

simple ion - an atom that has lost or gained one or more electrons to conform with the octet rule. All ions have a charge.

metal - an element that can form positive ions. Metals are lustrous, malleable, ductile, and conduct heat and electricity well. Generally, metals are shown on the left side of the staircase line on the periodic table

nonmetal - an element that can form negative ions. Nonmetals are dull in appearance, brittle and conduct heat and electricity very poorly.

ionic compound - a compound that is composed of positive and negative ions. Ionic compounds usually contain metallic and nonmetallic ions.

molecule - the smallest particle that exists in a free state. A molecule is a group of chemically bonded atoms or a single atom as for the noble gases.

element - a pure substance which contains only one kind of atoms

compound - a pure substance which contains more than one kind of atom bonded together. These atoms can not be physically separated from one another.

molecular compounds - compounds that exist only as molecules. Molecular compounds usually contain only nonmetallic elements.

acid - an aqueous solution of a compound containing hydrogen which conducts electricity, tastes sour, turns blue litmus red, neutralizes bases and reacts with active metals to form hydrogen gas. The term acid sometimes also refers to the compound dissolved. (Refer to Unit N, *Acids and Bases.*)

base - an aqueous solution that reacts with acids. Bases turn litmus blue, taste bitter, and conduct electricity. (Refer to Unit N, *Acids and Bases.*)

salt - any ionic compound that could be produced by an acid-base reaction

complex ion - a stable group of atoms that have lost or gained one or more electrons. The atoms in the complex ion are covalently bonded.

REVIEW OF CHEMISTRY 10 & 20

ELEMENTS AND PERIODIC TABLE (REMEDIAL)

	Species Name	Species Symbol	Number Protons	Number Electrons	Common Use	Net Charge	Electrons in Each Energy Level			# Valence Electrons
							1	2	3	
	Example: carbon atom	C	6	6	main constituent of coal	0	2	4	0	4
1.		Al			cookware, house siding					
2.	chloride ion				production of $Cl_2(g)$ by electrolysis					
3.		Si			transistors and semi-conductor diodes					
4.		Mg^{2+}			causes water hardness					
5.	neon atom				neon lights					
6.					in table salt	1^+	2	8		
7.			16		production of sulfuric acid					6
8.			16	18	in tarnish of silver					
9.					80% of air consists of pairs of these atoms	0	2	5	0	
10.		H^+			provided by all acis					

TABLE K1

REVIEW OF CHEMISTRY 10 & 20

IONIC COMPOUNDS, ACIDS, MOLECULAR COMPOUNDS

	Ionic Compounds	Acids	Molecular Compounds
Formula of Pure Substance	empirical formula (indicates simplest ratio of ions)	molecular formula	molecular formula (indicates actual number of atoms in a molecule)
Type of Bond and Arrangement of Electrons	ionic (electrovalent) bond (atoms gain or loose electrons to form ions)	covalent bond	covalent bond (atoms share pairs of electrons)
Melting Point and State of Room Temperature	high melting point; solid at room temperature	low melting point; solid, liquid or gas at room temperature	low melting point; solid, liquid or gas at room temperature
Color of Aqueous Solution (if substance is soluble)	colorless or colored	colorless or colored	colorless only
Electrical Conductivity of Aqueous Solution (if substance is soluble)	conducts electricity	conducts electricity	does not conduct electricity
Nomenclature	ionic nomenclature (Stock or Classical system)	ionic nomenclature or acid nomenclature	molecular nomenclature (prefix system)
Examples	$NaCl_{(s)}$, $NaOH_{(s)}$, $CuSO_{4(s)}$, $SnF_{2(s)}$	$HCl_{(g)}$ . . . . $HCl_{(aq)}$ $H_2SO_{4(\ell)}$ . . . $H_2SO_{4(aq)}$ $C_6H_5COOH_{(s)}$ . . $C_6H_5COOH_{(aq)}$	$NH_{3(g)}$ $CH_{4(g)}$ $CCl_{4(\ell)}$ $C_2H_5OH_{(\ell)}$ $P_4O_{10(s)}$ $C_{12}H_{22}O_{11(s)}$

REVIEW OF CHEMISTRY 10 & 20

NOMENCLATURE

Polyatomic Molecular Elements

In naming elements it is *not* necessary to indicate the number of atoms that are bonded together in a molecule; *e.g.*, hydrogen is the name of H . The polyatomic molecular elements are: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2, S_8 (crystalline) and P_4 (white phosphorus).

Nomenclature for Molecular Compounds-Prefix System

Name the compound using the Greek or Roman (nona) prefix to indicate the number of each kind of atom covalently bonded to one another. The prefixes should be learned.

Chemical Prefixes	
mono = 1	hexa = 6
di = 2	hepta = 7
tri = 3	octa = 8
tetra = 4	nona = 9
penta = 5	deca = 10

Formulas to be Learned	
Formula	Name
$O_{3(g)}$	ozone
HOH (H_2O)$_{(\ell)}$	water
$NH_{3(g)}$	ammonia
$CH_{4(g)}$	methane
$C_{12}H_{22}O_{11(g)}$	sucrose
$CH_3OH_{(\ell)}$	methanol
$C_2H_5OH_{(\ell)}$	ethanol

Alkanes to be Learned	
Formula (C_nH_{2n+2})	Name
CH_4	methane
C_2H_6	ethane
C_3H_8	propane
C_4H_{10}	butane
C_5H_{12}	pentane
C_6H_{14}	hexane
C_7H_{16}	heptane
C_8H_{18}	octane
C_9H_{20}	nonane
$C_{10}H_{22}$	decane

Naming Ionic Compounds

When naming any ionic compound the rule is to simply write the name of the cation (positive ion) followed by the name of the anion (negative ion). For ions with more than one charge the Stock System name (Roman numerals) is preferred over the classical system name (ic, ous).

Acids

Acids are named by changing the hydrogen compound name to the acid name using the rules given on the *ALCHEM* periodic table.

Hydrated Compounds

The complete formula of a hydrated compound is written by following the formula of compound with a dot and the proper number of water molecules. In the full name the water is read as *hydrate* using the usual chemistry prefixes.

E.g., $CuSO_4 \cdot 5H_2O$ - copper(II) sulfate pentahydrate

REVIEW OF CHEMISTRY 10 & 20

NOMENCLATURE REMEDIAL

	i, m, or a	Chemical Formula	Common Name and/or Use	Chemical Name of Substance
			rust remover	phosphoric acid
		$NaNO_2(s)$	meat preserver	
		$NO_2(g)$	component of smog	
		$HCl(aq)$	very common acid; cleans concrete & galvanized iron; also gastric acid	
		$HCl(g)$	used to make hydrochloric acid	
			bluestone fungicide	copper(II) sulfate pentahydrate
			used for stating P-equivalent in fertilizer	solid diphosphorus pentaoxide
			found in sour natural gas	hydrosulfuic acid
		$H_2(g)$	hydrogen torch welding, weather balloons	
			ingredient of fertilizer	ammonium phosphate
		$H_2O_2(aq)$	disinfectant, mouthwash, bleach	
			household cleaner	aqueous ammonia
			used to make fertilizers	ammonia gas
		$K_2SO_3(s)$	skin fungicide, treatment of ringworm	
		$CH_3OH(\ell)$	*wood alcohol*; windshield washer fluid, duplicating fluid	
2.			prevents tooth decay	tin(II) fluoride stannous fluoride
3.		$Na_2SO_4 \cdot 10H_2O(s)$	Glauber's salt, medicine	

REVIEW OF CHEMISTRY 10 & 20

NOMENCLATURE REMEDIAL

Provide either the chemical name or formula as required. With the chemical formula indicate the physical state of the substance. Classify the chemicals as either an ionic (i) or molecular (m) substance or an acid (a).

	i, m, or a	Chemical Formula	Common Name and/or Use	Chemical Name of Substance
	e.g. i	$NaHSO_{4(s)}$	Saniflush, bowl cleaner	sodium bisulfate sodium hydrogen sulfate
1.		$H_2SO_{4(aq)}$	the most used industrial acid	
2.			toxic cleaning solvent	carbon tetrachloride
3.		$HgS_{(s)}$	cinnabar, a mercury ore	
4.		$NaHCO_{3(s)}$	baking soda	
5.			drinking alcohol	ethanol
6.			epsom salts medicine	magnesium sulfate heptahydrate
7.			milk of magnesia, stomach antacid, laxative	magnesium hydroxide
8.		$Na_2S_2O_3 \cdot 5H_2O_{(s)}$	photographic hypo, antichlor	
9.		$O_{2(g)}$	the gas that sustains respiration and combustion	
10.		$NaHSO_{3(s)}$	yeast inhibitor	
11.			siderite, iron ore	iron(II) carbonate
12.		$Fe_2O_{3(s)}$	hematite, iron ore	
13.			most common sweetener	sucrose
14.		$CaSO_4 \cdot 2H_2O_{(s)}$	gypsum, building material	
15.			water glass; preserves eggs, adhesive cement, fireproofing	sodium silicate
16.			bleaching of paper, water treatment	chlorine gas

REVIEW OF CHEMISTRY 10 & 20
SIGNIFICANT DIGITS, MOLES AND MOLARITY

Definition

Significant digits are all of the certain digits from a measurement plus one uncertain (estimated) digit.

Counting Significant Digits

1. Count all digits from 1 to 9 plus zeros in between and following other digits.
2. Do not count zeros in front of a value because they only serve to set the decimal place.

Exact numbers

Exact numbers are not uncertain and are said to have an infinite number of significant digits.

1. Numbers that are defined; *e.g.*, 1000 kg = 1 t.
2. Numbers that result from counting objects; *e.g.*, 32 students.

The Addition and Subtraction Rule

Round-off all the values to the *least number of decimal places* and then add or subtract.

The Multiplication and Division Rule

Multiple or divide and then round-off to the *least number of significant digits.*

Definitions

Mole - the amount of substance in a system which contains as many elementary entities as there are atoms in exactly 0.012 kg of carbon 12. (SI definition) Note: Entities can be atoms, molecules, electrons, ions, formula units or other particles. Carbon-12 is the most common isotope of carbon-containing 6 protons, 6 neutrons, and 6 electrons. To three significant digits the number of atoms in 0.012 kg is 6.02×10^{23} (Avogadro's number). The mole can be thought of as 6.02×10^{23} particles.

Molar mass - the mass of one mole of substance is units of grams per mole (g/mol)

Molar concentration - a measure of the concentration of solutions expressed in units of moles of solute per litre of solution-symbolized as mol/ℓ or M. The units M are equivalent to the units mol/ℓ and is read as *molar*. The molar concentration of a solution is often referred to as the *molarity* of a solution.

Formulas

$$\boxed{\# \text{ moles} = \frac{\text{mass}}{\text{molar mass}}} \longrightarrow \boxed{\text{mass} = \# \text{ moles} \times \text{molar mass}}$$

$$\boxed{M = \frac{\# \text{ moles}}{\# \text{ litres}}} \longrightarrow \boxed{\# \text{ moles} = M \times \# \text{ litres}} \longrightarrow \boxed{\# \text{ litres} = \frac{\# \text{ moles}}{M}}$$

REVIEW OF CHEMISTRY 10 & 20

THE MOLE AND MOLAR CONCENTRATION (REMEDIAL)

	Name, Formula and Use	Molar Mass	Mass	# Moles	Molarity	Solution Volume
e.g.	$Cu(NO_3)_2$ copper(II) nitrate	1 Cu = 1 x 63.5 = 63.5 2 N = 2 x 14.0 = 28.0 6 O = 6 x 16.0 = 96.0 187.5 or 188 g/mol	226 g (given)	$\# \text{ moles} = \frac{\text{mass}}{\text{molar mass}}$ $Cu(NO_3)_2 = \frac{226 \text{ g}}{187.5 \text{ g/mol}}$ = 1.21 mol	$M = \frac{\# \text{ mole}}{\# \text{ litres}}$ $Cu(NO_3)_2 = \frac{1.20 \text{ mol}}{4.00}$ = 3.000 M	4.00 ℓ (given)
1.	carbon tetrachloride (common solvent dry cleaning)		16.9 g			
2.	SiO_2 (sand)			0.45 mol		
3.	sulfuric acid (most common industrial acid)			1.40 mol	0.400 M	

REVIEW OF CHEMISTRY 10 & 20

THE MOLE AND MOLAR CONCENTRATION (REMEDIAL)

	Name, Formula and Use	Molar Mass	Mass	# Moles	Molarity	Solution Volume
4.	NaCl (table salt)				0.0600 M	70.0 ml
5.	$NaHSO_4$ (Sani-Flush)			0.0250 mol	0.400 M	
6.	potassium dichromate (a common oxidizing agent)		35.0 g			200 ml
7.	$Pb(CH_3COO)_2 \cdot 3H_2O$ (a soluble lead salt)		13.0 g		0.0320 M	

REVIEW OF CHEMISTRY 10 & 20
PREPARATION AND DILUTION OF SOLUTIONS

Preparation of Solutions

Dissolve the correct mass of solute in less than the required amount of solvent. Then add solvent to bring the total volume to the desired value.

Dilution of Solutions

When a solution is diluted only the amount of solvent is increased. Therefore, the number of moles of solute in the *initial* concentrated solution is equal to the number of moles of solute in the *final* diluted solution.

Since # moles = Mx # litres,

$$M_i \times \# \text{ litres}_i = M_f \times \# \text{ litres}_f$$

i = initial

f = final

Definitions

Solubility - the concentration of solute in a saturated solution at a given temperature. The units of solubility are usually grams of solute per 100 ml of solution.

Molar Solubility - the molar concentration of a saturated solution at a given temperature.

REVIEW OF CHEMISTRY 10 & 20

SOLUTIONS AND SOLUBILITY

1. The solubility of sodium chloride (table salt) is 36.5 g per 100 ml of solution at room temperature. Calculate the molar solubility of sodium chloride.

2. A saturated solution of sodium hydroxide at room temperature has a concentration of 19.1 M. What mass of sodium hydroxide does a 100 ml sample of saturated solution contain?

3. What volume of 6.00 M sulfuric acid is required to make 200 ml of 0.75 M acid solution?

4. Describe how to prepare a 0.100 M solution of copper(II) sulfate pentahydrate? (Show calculations and steps.)

REVIEW OF CHEMISTRY 10 & 20

CHEMICAL REACTIONS

Chemical Reaction Clues

1. F[illegible]mation of a precipitate
2. [illegible]ion of a gas
3. Color change
4. Energy changd

Conservation Laws

In a chemical reaction all of the following are conserved.

1. the number of each kind of atom
2. mass
3. energy

Information in a Balanced Chemical Equation

A balanced chemical equation shows the chemical composition, state, and mole relations of all substances involved as well as the reaction type.

Classification of Chemical Reactions by the Chemical Composition of the Reagents Involved

The product of a chemical reaction can be predicted from the type of chemical reaction.

1. *Simple composition* (sc): element + element ⟶ compound
2. *Simple decomposition* (sd): compound ⟶ element + element
3. *Single replacement* (sr): element + compound ⟶ element + compound
4. *Double replacement* (dr): compound + compound ⟶ compound + compound
5. *Hydrocarbon combustion* (hc): hydrocarbon + oxygen ⟶ carbon dioxide + water
6. *Other* (o): Reaction cannot be specified as being any of the preceding five types.

Classification of Reactions by Energy Changes

Endothermic reaction - energy enters the system;
i.e., energy is absorbed by the system; *i.e.*, energy is added to the system.
For example, in electrolysis energy is consumed (absorbed).

$$2H_2O_{(\ell)} + \text{ energy} \longrightarrow 2H_{2(g)} + O_{2(g)}$$

Exothermic reaction - energy leaves the system;
i.e., energy is given off by the system; *i.e.*, energy is removed from the system.
For example, in hydrogen welding heat is released.

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)} + \text{ energy}$$

REVIEW OF CHEMISTRY 10 & 20

CHEMICAL REACTIONS

Using the simplest whole number coefficients write a balanced chemical equation for each of the following chemical reactions. Include the state of matter as subscripts in the balanced equation. Identify the reaction type in the space provided.

Example:

In water treatment plants *hardness* in the form of calcium sulfate is removed from the water by adding a washing soda solution.

Reaction type: *double replacement*

Balanced equation: $CaSO_{4(aq)} + Na_2CO_{3(aq)} \longrightarrow CaCO_{3(s)} + Na_2SO_{4(aq)}$

Word equation: calcium sulfate + sodium carbonate $\longrightarrow$ calcium carbonate + sodium sulfate

1. Combustion occurs in a photographic flashbulb.

   Reaction type:

   Balanced equation:

   Word equation:

2. Propane gas is burned in a camp stove.

   Reaction type:

   Balanced equation:

   Word equation:

3. Phosphoric acid solution is used as rust (Fe_2O_3) remover.

   Reaction type:

   Balanced equation:

   Word equation:

4. The test for hydrogen is a *pop* when ignited in the presence of air.

   Reaction type:

   Balanced equation:

   Word equation:

5. Milk of magnesia (magnesium hydroxide) is used to counteract excess stomach acidity (hydrochloric acid).

   Reaction type:

   Balanced equation:

   Word equation:

CHEMICAL REACTIONS

6. The cap is removed from a carbonated beverage. The carbonic acid in the beverage decomposes to produce carbon dioxide and water.

   Reaction type:

   Balanced equation:

   Word equation:

7. Nickel is plated with copper using a copper(II) sulfate solution.

   Reaction type:

   Balanced equation:

   Word equation:

8. Baking soda (sodium hydrogen carbonate) is used to neutralize a car battery acid (sulfuric acid) spill.

   Reaction type:

   Balanced equation:

   Word equation:

9. Mercury(II) oxide is decomposed into its elements. (Joseph Priestley discovered oxygen using this reaction.)

   Reaction type:

   Balanced equation:

   Word equation:

10. Soap (sodium stearate) may react with hard water ($MgSO_{4(aq)}$) to precipitate bath-tub-ring.

    Reaction type:

    Balanced equation:

    Word equation:

11. Iron rusts in moist air (water and oxygen) to cause millions of dollars damage each year. (Rust is iron(III) oxide trihydrate.)

    Reaction type:

    Balanced equation:

    Word equation:

12. Sulfur dioxide emissions from industries combines with water vapor to produce *acid rain* (sulfurous acid).

    Reaction type:

    Balanced equation:

    Word equation:

REVIEW OF CHEMISTRY 10 & 20

EQUATIONS

Equations

Chemical equations are a shorthand way of representing what occurs in chemical reactions. The same notation can be extended to include changes of state and solvation (the process of dissolving).

Types of Solutes

nonelectrolytes - substances which dissolve to yield solutions that *do not* conduct electricity. Nonelectrolytes are usually molecular substances and in solution exist as molecules.

electrolytes - substances which dissolve to yield solutions that *do* conduct electricity. Electrolytes generally include ionic compounds and in solution exist as ions. (Acids are also electrolytes. See Unit N, *ALCHEM 30*.)

Remedial Problems

Write a balanced chemical equation for each of the following statements. Indicate the state of each species involved; *e.g.*, bluestone (copper(II) sulfate pentahydrate) is dissolved in water to make a fungicide spray for plum trees.

$$CuSO_4 \cdot 5H_2O_{(s)} \longrightarrow Cu^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

E.g., Baking soda (sodium bicarbonate) dissolves in water.

$$NaHCO_{3(s)} \longrightarrow Na^{+}{}_{(aq)} + HCO_3{}^{-}{}_{(aq)}$$

1. Methanol is evaporated.
2. Sodium hydroxide dissolves in water.
3. Ethanol dissolves in water.
4. Ethanol is burned.
5. Bromine dissolves in ethanol.
6. Potassium burns in air.
7. Potassium reacts with water.
8. Potassium dissolves in mercury.
9. Iron(II) nitrate hexahydrate dissolves in water.
10. Epsom salts (magnesium sulfate heptahydrate) are dissolved in water.

Notes:

1. Solvents are indicated as subscripts of products *not* as reagents.
2. Water of hydration is included in the water of solvation and is *not* written as a product.

REVIEW OF CHEMISTRY 10 & 20
IONIC CONCENTRATION

Calculating Ionic Concentrations

Step 1: Write the balanced equation for the dissociation.

Step 2: Use the formula,

$$\text{required concentration} = \text{given concentration} \times \frac{\text{coefficient of required species}}{\text{coefficient of given species}}$$

Note: This formula applies only to solutions.

Example:

1. In a 0.23 M $Al_2(SO_4)_3$ solution, what is the concentration of each ion?

   Step 1: Write a balanced equation. $Al_2(SO_4)_{3(s)} \longrightarrow 2Al^{3+}_{(aq)} + 3SO_4^{2-}{}_{(aq)}$

   Step 2: Calculate the molar concentration of each ion using the mole ratio.

   $$\text{required concentration} = \text{given concentration} \times \frac{\text{coefficient of required species}}{\text{coefficient of given species}}$$

   $$M_{Al^{3+}} = 0.23\ M \times \frac{2}{1} = 0.46\ M \qquad M_{SO_4^{2-}} = 0.23\ M \times \frac{3}{1} = 0.69\ M$$

   The concentration of the Al^{3+} ion is 0.46 M and the concentration of SO_4^{2-} ion is 0.69 M.

2. What amount of ammonium sulfate is required if 100 ml of solution with a 0.100 M $NH_4^+{}_{(aq)}$ ion concentration is to be made?

   Step 1: Write a balanced equation. $(NH_4)_2SO_{4(s)} \longrightarrow 2NH_4^+{}_{(aq)} + SO_4^{2-}{}_{(aq)}$

   Step 2: Calculate the concentration of $(NH_4)_2SO_{4(aq)}$.

   $$\text{required concentration} = \text{given concentration} \times \frac{\text{coefficient of required species}}{\text{coefficient of given species}}$$

   $$M_{(NH_4)_2SO_4} = 0.100\ M \times \frac{1}{2} = 0.0500\ M$$

   Step 3: Calculate the amount of $(NH_4)_2SO_4$ required

   $$\#\text{moles}_{(NH_4)_2SO_4} = M \times \#\text{litres} = 0.0500\ \text{mol}/\ell \times 0.100\ \ell = 0.00500\ \text{mol or } 5.00 \times 10^{-3}\ \text{mol.}$$

   The amount of ammonium sulfate required is 5.00×10^{-3} mol.

Remedial Problems

Find the molar concentration of each ion in the following solutions. Work does not have to be shown.

1. 0.120 M Na_2CO_3 (washing soda) ____________
2. saturated table salt solution (5.3 M) ____________
3. 0.0621 M ammonium dichromate ____________
4. 0.84 M Na_3PO_4 (commercial name is TSP or trisodium phosphate) ____________

Find the concentration if the concentration of the ion is given.

5. 0.500 M $Na^+_{(aq)}$ in a sodium carbonate solution. ____________
6. 0.20 M $Cr_2O_7^{2-}{}_{(aq)}$ in a potassium dichromate solution. ____________
7. 0.575 M $Cl^-_{(aq)}$ in a iron(III) chloride solution. ____________
8. 0.00233 M $CH_3COO^-_{(aq)}$ in a lead(II) acetate solution. ____________

REVIEW OF CHEMISTRY 10 & 20

REACTIONS IN SOLUTION

Types of Reactions in Solution

Three common types of reactions in solution are single replacement reactions, precipitation reactions and neutralizing reactions. The latter two represent different types of double replacement reactions.

1. Single Replacement Reactions in Solution

Single replacement reactions usually take place when a sample of an element is added to an aqueous solution of a compound. The element reacts with a compound to form a new element and a new compound. The general equation for a single replacement reaction is

$$A + BC \longrightarrow B + AC$$

a. Metals replace metals. Nonmetals replace nonmetals.

active metal + acid ⟶ hydrogen + salt

$$2Al_{(s)} + 3H_2SO_{4(aq)} \longrightarrow 3H_{2(g)} + Al_2(SO_4)_{3(aq)}$$

b. active metal + water ⟶ hydrogen + a hydroxide

$$2Na_{(s)} + HOH_{(\ell)} \longrightarrow H_{2(g)} + 2NaOH_{(aq)}$$

c. metal + salt ⟶ metal + salt

$$Zn_{(s)} + CuSO_{4(aq)} \longrightarrow Cu_{(s)} + ZnSO_{4(aq)}$$

d. nonmetal + salt ⟶ nonmetal + salt

$$Cl_{2(g)} + 2NaBr_{(aq)} \longrightarrow Br_{2(\ell)} + 2NaCl_{(aq)}$$

2. Double Replacement Reactions in Solution

a. Precipitation Reactions

Precipitation reactions occur when two compounds in solution react to produce a new compound that has low solubility. The new compound of low solubility is called a precipitate. Precipitation is identified in an equation by a product with the subscript "s ". For example:

$$BaCl_{2(aq)} + Na_2SO_{4(aq)} \longrightarrow 2NaC\ell_{(aq)} + BaSO_{4(s)}$$

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \longrightarrow \text{[illegible]}_{(aq)} + \text{[illegible]}_{(s)}$$

Solid particles of precipitate settle out and can be separated from the remaining solution by filtration. The liquid which goes through the filter paper is called the *filtrate*. The solid precipitate caught in the filter paper can be dried and its mass determined.

The rules for determining the solubility of ionic compounds are given on the ALCHEM data sheet.

b. Neutralization Reactions

Neutralization reactions are reactions between acids and bases in solution. The products of a neutralization are water and salt.

acid + base ⟶ water + salt

For example: $HCl_{(aq)} + LiOH_{(aq)} \longrightarrow HOH_{(\ell)} + LiCl_{(aq)}$

NET IONIC EQUATIONS

Writing Ionic Equations

Many chemical reactions can be represented by three different kinds of equations: *nonionic equations*, *total ionic equations* and *net ionic equations*. For reactions in aqueous solution, the most correct are ionic equations since in the ionizing water media, substances that are electrolytes undergo dissociation into ions. The ionic species in aqueous solution subsequently react as ions.

1. Nonionic Equations

In nonionic equations, the elements and compounds are written in their molecular or formula unit forms.

Example:

$$2AgNO_{3(aq)} + BaCl_{2(aq)} \longrightarrow 2AgCl_{(s)} + Ba(NO_3)_{2(aq)}$$

2. Total Ionic Equations

In total ionic equations, elements and compounds are written in the forms in which they are predominately present: electrolytes as ions; nonelectrolytes, precipitates and gases in their molecular or formula unit forms.

Example:

$$2Ag^{1+}_{(aq)} + 2NO^{1-}_{3(aq)} + Ba^{2+}_{(aq)} + 2Cl^{1-}_{(aq)} \longrightarrow 2AgCl_{(s)} + Ba^{2+}_{(aq)} + 2NO^{1-}_{3(aq)}$$

3. Net Ionic Equations

In net ionic equations, only those molecules, formula units or ions that have changed (predominant reacting species) are included in the equation; ions or molecules that do not change (spectator species) are omitted.

Example:

$$Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \longrightarrow AgCl_{(s)}$$

The following is a summary of rules to observe when writing ionic equations.

1. Electrolytes are written in their ion form.
2. Nonelectrolytes are written in their molecular forms.
3. The net ionic equation includes only those substances that have undergone a chemical change; i.e., predominant reacting species.
4. Insoluble substances, precipitates and gases are written in their molecular or formula unit forms.
5. Equations must be balanced, both in atoms and in electrical charge.

Example:

Write the nonionic, total ionic and net ionic equations representing the reactions between aqueous sodium sulfate and aqueous barium nitrate.

$$Na_2SO_{4(aq)} + Ba(NO_3)_{2(aq)} \longrightarrow BaSO_{4(s)} + 2NaNO_{3(aq)} \quad \text{(nonionic eq.)}$$

$$2Na^{+}_{(aq)} + SO^{2-}_{4(aq)} + Ba^{2+}_{(aq)} + 2NO^{-}_{3(aq)} \longrightarrow BaSO_{4(s)} + 2Na^{+}_{(aq)} + 2NO^{-}_{3(aq)} \quad \text{(total ionic eq.)}$$

$$SO^{2-}_{4(aq)} + Ba^{2+}_{(aq)} \longrightarrow BaSO_{4(s)} \quad \text{(net ionic eq.)}$$

REVIEW OF CHEMISTRY 10 & 20

NONIONIC, TOTAL IONIC AND NET IONIC EQUATION

Write the nonionic equation, the total ionic equation and the net equation for each of the following reactions. Indicate the state of each species in each reaction.

1. A solution of iron(III) chloride is tested for the presence of the iron(III) ion by the addition of dilute sodium hydroxide. A red precipitate indicates a positive result.

2. Aqueous solutions of potassium sulfate and barium bromide are mixed.

3. An aqueous solution of washing soda (Na_2CO_3) is added to hard water containing magnesium sulfate. The magnesium ions are precipitated from water.

4. A copper(II) sulfate solution is tested for the presence of copper(II) ion by the addition of an aqueous solution of sodium sulfide. A black precipitate indicates a positive result.

5. Liquid bromine is added to an aqueous solution of sodium iodide. The solution turns brown indicating the presence of aqueous molecular iodine.

NONIONIC, TOTAL IONIC AND NET IONIC EQUATION

6. Zinc can be shown to be a chemically more reactive metal than copper by reacting zinc with aqueous copper(II) sulfate.

7. Potassium metal reacts violently with water.

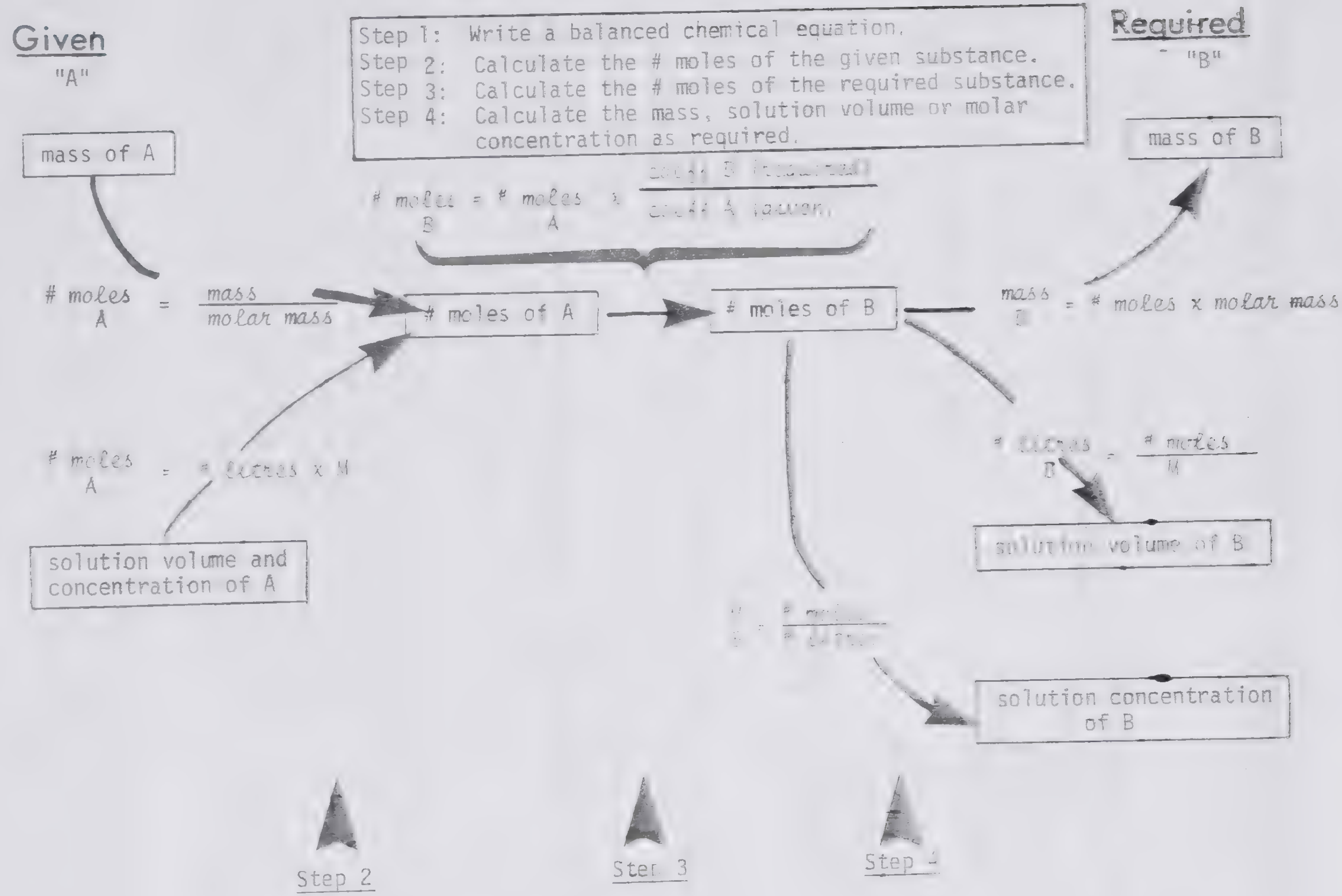
8. Nickel is extracted from aqueous nickel(II) sulfate by the addition of copper.

9. An aqueous solution of silver nitrate produces a white precipitate when reacted with aqueous sodium chloride. The white precipitate indicates the presence of silver ions in the original solution.

10. Chlorine water (aqueous molecular chlorine) is mixed with a solution of potassium bromide. The solution turns orange indicating the presence of aqueous molecular bromine.

REVIEW OF CHEMISTRY 10 & 20

STOICHIOMETRY FLOWCHART



Figure

REVIEW OF CHEMISTRY 10 & 20

STOICHIOMETRY LAB - LAB K1

Purpose:

To quantitatively determine the concentration of sulfate ions in a sample of sodium sulfate of unknown concentration using a 0.0500 M solution of $BaCl_{2(aq)}$.

Materials:

Day I:

150 ml of 0.0500 M $BaCl_2$

10.0 ml of $Na_2SO_{4(aq)}$ of unknown concentration

2 - 250 beakers
1 - stirring rod
1 - 10 ml pipet & bulb
1 - watch glass

Day II:

1 - filter paper
1 - filter funnel
1 - funnel rack
1 - ring stand
1 - stirring rod c/w rubber policeman
1 - 50 ml or 100 ml graduated cylinder

Prelab Exercise:

Write the balanced chemical equations for the reaction in this lab.

nonionic equation: ______________________________

total ionic equation: ______________________________

net ionic equation: ______________________________

Procedure:

Day I:

1. Use a graduated cylinder to obtain about 25 ml of sodium sulfate solution of unknown concentration. Pipet 10.0 ml of the sodium sulfate solution into a 250 ml beaker.
2. Use a 250 ml beaker to obtain about 150 ml of 0.0500 M Barium chloride solution. all but about 25 ml of the 150 ml of 0.0500 M barium chloride solution to the sodium sulfate solution. Stir for two or three minutes and let stand for about five minutes.
3. Add another 25 ml of 0.0500 M barium chloride solution to verify that all the sulfate has precipitated. (The barium chloride solution is in excess and the volume used will not be part of the calculations.).
4. Cover the beaker with a watch glass and let the mixture set overnight to get larger crystals.

Day II:

5. Determine and record the mass of a piece of filter paper to 0.01 g.
6. Filter the solution from Day I, Step 4. Make sure all of the precipitate has been transferred to the filter paper. Use water from a wash bottle and a stirring rod with a rubber policeman to remove any solid that sticks to the beaker. Wash the precipitate with distilled water.
7. Carefully remove the filter paper containing the barium sulfate from the funnel and place it unfolded onto a watchglass. Label the watchglass and place it in a fume hood to dry.

STOICHIOMETRY LAB - LAB K.1

Day III:

8. Carefully remove the filter paper containing the precipitate from the watchglass. Determine and record the mass of the filter paper plus precipitate.

Observations:

Mass of filter paper plus barium sulfate ____________________

Mass of filter paper ____________________

Calculations:

1. Mass of barium sulfate precipitate ____________________

2. Calculate the molar concentration of the sulfate ions in the original sample.

In industry sulfate ion concentrations are reported in units of *grams of sulfur per litre of solution*. To convert the results of this lab to this form find:

3. the mass of sulfate ions in 1.000 ℓ of solution.

4. the percentage of sulfur in sulfate ions.

5. the mass of sulfur per litre of solution.

REVIEW OF CHEMISTRY 10 & 20

STOICHIOMETRY LAB - LAB K1

Question:

Chemical plants are limited in the amount of sulfate ions they are allowed to release into the environment. The effluents of each plant is monitored to see that the standards set by the government are not exceeded. The test used involves the same chemical reaction as the one used in this experiment.

The effluent of a chemical plant was tested for sulfur in the form of sulfates. It was found that a 100 ml sample produced 0.04200 g of barium sulfate when reacted with a solution of barium chloride. Calculate the concentration of sulfate ions in the effluent and the number of grams of sulfur per litre in the effluent.

REVIEW OF CHEMISTRY 10 & 20

STOICHIOMETRY

1. How many moles of oxygen are needed to completely oxidize 3.60 g of steel wool (assume pure iron) to produce iron(III) oxide?

2. What mass of water would be produced from the complete combustion of 15.0 g of C_8H_{18} (one of the components of gasoline)?

3. To what final volume must 45.0 ml of 6.00 M $HCl_{(aq)}$ be diluted to obtain a 0.400 M solution?

STOICHIOMETRY

4. What mass of mercury(II) oxide must be decomposed to yield 5.50 g of mercury?

5. An excess of a barium chloride solution was added to 400 ml of silver nitrate solution and 5.10 g of precipitate was formed. What was the concentration of the silver nitrate solution?

6. When 25.0 ml of a 0.400 M hydrochloric acid solution was pipetted into an Erlenmeyer flask and titrated with a solution of potassium hydroxide the following data were obtained. Calculate the concentration of the potassium hydroxide solution.

Buret Reading (KOH)	Trial 1	Trial 2	Trial 3
Final	29.8 ml	39.8 ml	49.6 ml
Initial	20.2 ml	29.8 ml	39.8 ml
Volume Used			

REVIEW OF CHEMISTRY 10 & 20

STOICHIOMETRY

7. What volume of 0.500 M copper(II) sulfate solution must be reacted with an excess of zinc to obtain a 1.00 g deposit of copper?

8. What mass of zinc will react with 50.0 ml of 6.00 M hydrochloric acid?

9. A 20.0 ml sample of iron(III) chloride solution is reacted with an excess of 1.00 M sodium hydroxide solution and 2.36 g of precipitate was obtained. What was the concentration of the iron(III) chloride solution?

STOICHIOMETRY

10. Describe in detail how you would prepare 1000 ml of a 0.400 M solution of corn sugar ($C_6H_{12}O_6 \cdot H_2O$).

11. What is the concentration of a sodium acetate solution made by diluting 200 ml of a 0.250 M solution to 5.000 ℓ?

12. If 20.5 g of $Ca(NO_3)_2$ is dissolved to form 200 ml of solution, what is the concentration of the anions in the solution?

REVIEW OF CHEMISTRY 10 & 20

LEWIS DIAGRAMS AND VSEPR THEORY

Definitions

chemical bond - a force of attraction between chemical species (atoms, ions, molecules)

covalent bond - a bond in which electron pairs are *shared* between atoms

ionic bond - the bond due to the electrostatic attraction between a positive and a negative ion. For ions to be formed one or more electrons must be transferred.

octet rule - an atom or ion with eight valence electrons is stable

Lewis Diagram of an atom - a diagram that represents the valence electrons of an atom as dots around the chemical symbol of the element

Lewis Diagram of a molecule - a diagram that shows the valence electrons of all the atoms in a molecule so arranged that the octet rule is satisfied for each atom. Electrons are always shown in pairs.

structural formula - a diagram of a molecule that represents atoms by the atomic symbol and electron pairs forming covalent bonds by lines

lone electron pair - an electron pair not involved in bonding

bonding electron pair - an electron pair shared by two atoms

VSEPR THEORY AND SHAPES OF MOLECULES

All discrete molecules have a definite three-dimensional shape. The Valence Shell Electron Pair Repulsion Theory (VSEPR theory) provides a relatively simple and reliable basis for understanding and predicting molecular geometry. The theory only requires that the number of valence electrons be known. Molecular geometry can be determined from Lewis electron-dot diagrams and VSEPR theory.

The VSEPR Theory proposes a set of rules for predicting molecular geometry based on the idea that, the arrangement in space of the covalent bonds formed by an atom depends on the *arrangement of electron pairs in the outermost on valence shell of the central atom.* For this course, the most relevant of the rules are:

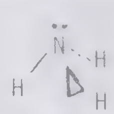
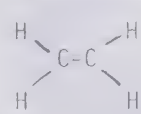
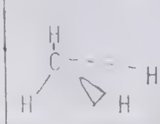
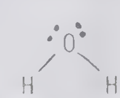
1. Valence electron pairs, both shared (bonded) and lone (nonbonded) arrange themselves around the central atom in a molecule in such a way as to minimize repulsion. Thus, the bonding and lone pairs of electrons take-up positions around the central atom as far away from one another as possible.

2. For purposes of predicting molecular geometry, treat multiple bonds (double and triple bonds) as one bond.

Table K2
Shapes Around the Central Atom

Shape	Number of Bonds Around the Central Atom	Number of Lone Electron Pairs Around the Central Atom
linear	2	0
trigonal planar	3	0
tetrahedral	4	0
pyramidal	3	1
V-shaped	2	2

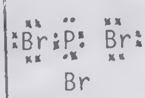
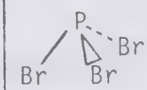
Table K3
Examples Showing 3-Dimensional Representation

O=C=O				
linear about the carbon atom	pyramidal about the nitrogen atom	trigonal planar about the carbon atom	tetrahedral about the carbon atom	V-shaped about the oxygen atom

REVIEW OF CHEMISTRY 10 & 20

SHAPES OF MOLECULES

Complete the following table.

	Formula of Compound	Lewis Diagram	Number of Lone Pairs on the Central Atom	Number of Bonds on the Central Atom	Name of Shape	Dimensional Representation
e.g.	PBr_3		1	3	pyramidal	
1.	H_2Se					
2.	C_2F_4					
3.	SiH_4					
4.	C_2H_2					
5.	H_2O_2					
6.	C_2H_6					
7.	HCHO					

POLARITY OF BONDS AND MOLECULES

Definitions

electronegativity - a measure of the relative attraction of an atom for its valence electrons. The electronegativity of each element is given on the *ALCHEM* periodic table.

nonpolar covalent bond - a covalent bond formed by two atoms of equal electronegativity. The electron pair (or pairs) is shared equally.

polar covalent bond - a covalent bond formed by two atoms of unequal electronegativity. The electron pair (pairs) is not shared equally.

dipole - the unequal distribution of electric charge

bond dipole - the unequal distribution of charge in a polar covalent bond. The atom with the higher electronegativity attracts the bonding electrons more strongly making that atom slightly negative. The other atom is then slightly positive.

nonpolar molecule - a molecule in which all the bond dipoles add to zero.

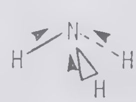
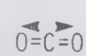
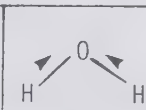
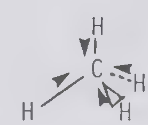
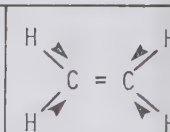
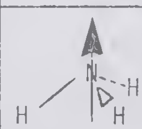
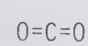
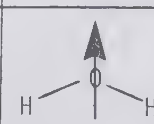
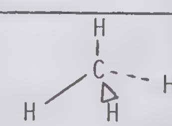
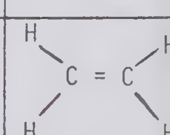
polar molecule - a molecule in which all the bond dipoles do not add to zero.

molecular dipole - the net dipole bound by adding all the bond dipoles.

Dipole Representation

A dipole is represented by an arrow next to the bond or an arrow through the molecule. The arrow points in the direction of the electron displacement; *i.e.* toward the more electronegative atom.

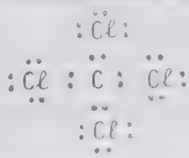
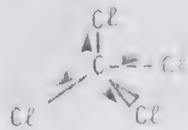
Table K4
Bond and Molecular Dipoles

3 Dimensional Representation Showing Bond Dipole					
3 Dimensional Representation Showing Molecular Dipole					
Addition of Bond Dipoles	do not cancel	cancel	do not cancel	cancel	cancel

REVIEW OF CHEMISTRY 10 & 20

POLARITY OF BONDS AND MOLECULES

Complete the following table: P-polar molecule NP-nonpolar molecule

	Formula	Lewis Diagram	Bond Dipole on 3-Dimensional Diagram	P or NP	Molecular Dipole on 3-Dimensional Diagram (if any)
e.g.	$C\,Cl_4$			NP	none
1.	$HOCl$				
2.	NBr_3				
3.	CH_3OH				
4.	$CHCl_3$				
5.	I_2				
6.	HCHO				

BONDING BETWEEN MOLECULES

Definitions

van der Waals forces - the forces of attraction between one molecule and another molecule. Van der Waals forces include London dispersion forces, dipole-induced-dipole interactions, and dipole-dipole interactions. The words *forces* and *interactions* are used because these forces are *weaker* than the ones referred to as *bonds*.

hydrogen bond - the attraction between the hydrogen atom (proton) bonded to a highly electronegative element and the lone pair electrons of a second highly electronegative atom. The hydrogen bond can be thought of as the interaction between two molecules or parts of a single molecule in which a proton is shared by two atoms. Only fluorine, chlorine, oxygen and nitrogen atoms have sufficiently high electronegativities to form hydrogen bonds.

Interactions of Nonpolar Molecules

In nonpolar substances the London dispersion forces are the van der Waals forces which are acting.

Interactions of Polar Molecules

In polar molecules the van der Waals forces consist of all three types of interactions. If the polar molecules contain a hydrogen atom bonded to a flourine, oxygen or nitrogen, hydrogen bonding is also present.

Factors Affecting the Strength of van der Waals Forces

1. The strength of the van der Waals forces increases with the total number of electrons in the molecules.

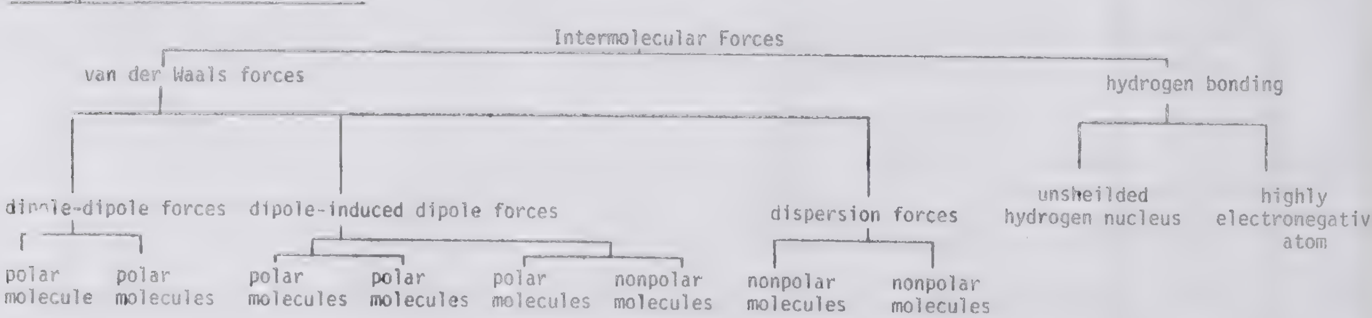
2. The strength of the van der Waals forces increases with the polarity of the molecules involved.

3. The strength of the van der Waal forces is influenced by the geometry of the molecules.

Relative Strength of Bonds

The approximate relationship of the bond strengths among van der Waals forces, hydrogen and covalent bonds is 1:10:100, respectively.

Summary of Intermolecular Forces



Intermolecular Forces and Polar Nature of Molecules
Figure K2

INTERMOLECULAR BONDING

Complete the following table.

Relation of Boiling Point to the Number of Electrons and to the Type of Intermolecular Forces

				Types of Intermolecular Forces			
				van der Waals			
Substance	Number of Electrons	Boiling Point (°C)		dipole-dipole	dipole-induced dipole	London dispersion	Hydrogen bonding
e.g., F_2	18	- 188				✓	
1. Cl_2		- 35					
2. Br_2		59					
3. I_2		184					
4. ClF		- 101					
5. BrF		- 20					
6. BrCl		5					
7. ICl		97					
8. IBr		116					
9. CH_4		- 162					
10. C_2H_6		- 87					
11. n - C_3H_8		- 45					
12. n - C_4H_{10}		- 0.5					
13. n - C_5H_{12}		36					
14. CF_4		- 129					
15. CCl_4		77					
16. CBr_4		189					
17. CH_3F		- 78					
18. CH_3Cl		- 24					
19. CH_3Br		3.6					
20. CH_3I		43					
21. CH_3OH		65					
22. C_2H_5F		- 38					
23. C_2H_5Cl		13					
24. C_2H_5Br		38					
25. C_2H_5I		72					
26. C_2H_5OH		78					

What relationship exists between number of electrons and types of intermolecular forces?

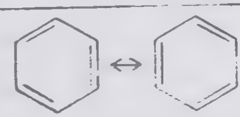
INTERMOLECULAR BONDING

Table K5

Summary of Chemical Bond Types

	Bond Type	Characteristics of Formulation	Some General Properties	Examples
Bonding Between Atoms	Covalent	Bonding electrons shared		
	a. nonpolar	Bonding electrons equally shared; uniform charge distribution	Low-melting solids, gases and liquids; nonconductors of electricity	H_2, N_2, O_2, Cl_2, I_2, P_2, S_8
	b. polar	Bonding electrons unequally shared; charge separation gives bond dipole	As above, but with slightly stronger bonds	HCl, H_2O, CH_4, CCl_4, NH_3
	c. covalent network	Bonding electrons equally shared to form three dimensional atomic lattice (crystal)	Very hard; very high melting point; insoluble in most ordinary solvents; nonconductors of electricity	C (diamond) SiO_2, SiC, BeO
	Ionic	Ion formation by electron transfer; electrostatic attraction between ions to form three dimensional ionic lattice (crystal)	Crystalline solids under ordinary conditions; high melting and boiling points; dissolve in polar liquids to form conducting solutions; electrical conductors in liquid phase	$NaCl$, $NaOH$, $CuSO_4$, NH_4Cl, $NaHCO_3$, $NaNO_3$
	Metallic	Positive ions in "sea" of mobile, delocalized valence electrons	Lustrous, malleable; good electrical conductors; wide range of melting points	Al, Fe, Cu, Zn, Ca, Na, Ag, Pb
Bonding Between Molecules	van der Waals	Weak electrostatic attraction among permanent temporary or induced molecular dipoles	Relatively low melting solids, gases or liquids because of relatively weak intermolecular forces	H_2, CO_2, He, H_2S, I_2, $C_{12}H_{22}O_{11}$, CCl_4, CH_4, S_8, HBr
	Hydrogen Bonding	Exposed hydrogen atom in one molecule attracted to negative end of highly electronegative atom in a polar molecule	Relatively high-melting solids, gases or liquids because of relatively strong intermolecular attraction	H_2O, HF, NH_3, C_2H_5OH

REVIEW OF CHEMISTRY 1 & 20
SUMMARY OF SOME ORGANIC COMPOUND GROUPS

GROUP	STRUCTURE (R REPRESENTS CARBON CHAIN)	NOMENCLATURE	SHAPE	a. PHYSICAL PROPERTIES b. CHEMICAL REACTIONS	USES AND OCCURRENCE
alkanes	[illegible]	— ane	tetrahedral around [illegible] atom	a. non polar; insoluble; low melting and boiling points b. substitution; combustion	fuels
alkenes	H or R - C = C - R orH	— ene	trigonal planar around each C [illegible]	a. same as alkanes b. addition	starting materials for many plastics
alkynes	H or R - C ≡ C - R orH	— yne	linear [illegible]	a. same as alkanes b. addition of one or two molecules of adding reagent	first member of the series used in oxyacetylene welding
aromatics		variable	each C [illegible] the ring has a [illegible] planar shape [illegible]. Molecule is [illegible]	a. non polar; insoluble in water b. substitution; no addition	very diverse--solvents, foods, drugs, explosives, mothballs
alcohols	R - OH	— anol	[illegible] atom	a. higher boiling; soluble because of hydrogen bonding b. many reactions, e. g., esterification, combustion	very diverse--antifreeze, alcoholic drinks, cosmetics, foods
acids	H or R - C - OH	— anoic acid	R - C(=O) - OH trigonal planar around C	a. high boiling; soluble due to hydrogen bonding b. all inorganic acid reactions: esterification	commonly occur in foods, waxes
esters	R - C(=O) - O - R′	— yl — anoate	R - C(=O) - O - R′ trigonal planar around C	a. insoluble in water b. can react with water to form an acid and alcohol	used as solvents and artificial flavors; commonly occur in animal fats and vegetable oils

REVIEW OF CHEMISTRY 10 & 20

ORGANIC CHEMISTRY

Rules for the IUPAC System of Nomenclature

1. Choose the longest continuous carbon chain that includes the
   a. functional group, or
   b. substitution, or
   c. side chain.

2. Number the carbon atoms in the longest continuous chain starting at the end closest to the
   a. functional group, or
   b. substitution, or
   c. side chain.

3. Give the compound the name of the alkane that has the same number of carbon atoms as the

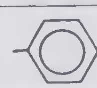
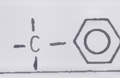
4. Change the suffix of the alkane to indicate the functional group present.

5. Locate the functional group by stating the number of the carbon atom to which the functional group is attached. Always start from the end of the longest chain that will locate the functional group using the lowest number possible.

6. Substitutions and side chains are identified by giving the name of each substitution and side chain. State the number of the carbon atom in the longest continuous chain to which the substitution or side chain is attached. Hydrocarbon side chains are named by changing the "[illegible]" suffix to "[illegible]". List the substitutions and side chains either
   a. alphabetically, or
   b. in order of size, or
   c. start at one end of the chain and proceed to the other end naming all successive substitutions and branches in order using the lowest numbers possible to locate all substitutions and side chains.

Table K7

Common Substitution Groups

Formula of Group	Prefix	Formula of Group	Prefix
-F	fluoro	$-NH_2$	amino-
-Cl	chloro-	$-CH_3$	methyl-
-Br	bromo-	$-CH_2CH_3$ or $-C_2H_5$	ethyl-
-I	iodo-	 or $-C_6H_5$	phenyl-
$-NO_2$	nitro-		benzyl-

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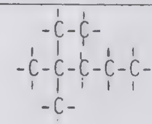
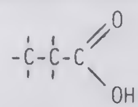
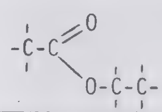
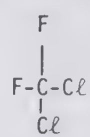
ORGANIC NOMENCLATURE (REMEDIAL)

	Systematic (IUPAC) Name	Structural Formula	Expanded Molecular Formula	Molecular Formula	Common Name and Use
eg.	2,2,3-trimethyl-pentane		$CH_3C(CH_3)_2CH(CH_3)CH_2CH_3$	C_8H_{18}	isoctane, the compound used to obtain the original octane rating of 100 when testing gasoline
eg.	propanoic acid		CH_3CH_2COOH	C_2H_5COOH	propionic acid, an emulsifier and mold inhibitar
eg.	ethyl ethanoate		$CH_3COOCH_2CH_3$	$CH_3COOC_2H_5$	banana flavor and odor
1.			CHCH		acetylene, welding
2.	1,2,3-propanetriol				glycerol or glycerine, cosmetics
3.					freon, refrigerant

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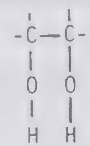
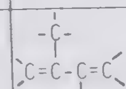
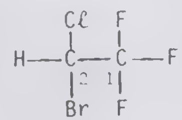
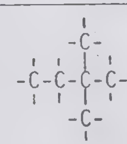
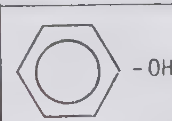
ORGANIC NOMENOLATURE (REMEDIAL)

	Systematic (IUPAC) Name	Structural Formula	Expanded Molecular Formula	Molecular Formula	Common Name and Use
4.			$CHCl_3$		chloroform, an anesthetic
5.					ethylene glycol, car radiator antifreeze
6.			$CH_3CH_2CH_2CH_2OH$		hydraulic fluid
7.	chloroethene				vinyl chloride, making of PVC plastics (polyvinyl chloride)
8.					isoprene, monomer natural rubber
9.	butanoic acid				butyric acid, odor of rancid butter, a very unpleasant odor

REVIEW OF CHEMISTRY 10 & 20

ORGANIC NOMENOLATURE (REMEDIAL)

	Systemic (IUPAC) Name	Structural Formula	Expanded Molecular Formula	Molecular Formula	Common Name and Use
10.					halothane, general inhalation anasthetic
11.			$CH_3CH_2CH_3$		fuel
12.	chloromethane				methyl chloride, used to make methanol
3.					neohexane, constituent of high test gasoline
14.	hexadecanoic acid		$CH_3(CH_2)_{14}COOH$		palmitic acid, in animal fat
15.	phenol				carbolic acid (archeic name), disinfectant

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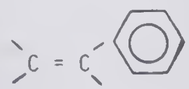
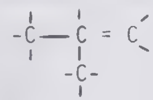
ORGANIC NOMENCLATURE (REMEDIAL)

	Systemic (IUPAC) Name	Structural Formula	Expanded Molecular Formula	Molecular Formula	Common Name and Use
16.	methylbutanoate				pineapple, odor and flavor
17.	(name as ethene compound)				styrene used to make polystyrene
18.			CH_3CHCH_2		propylene, making of the plastic poly propylene
19.					isobutylene, butyl rubber a soft plastic
20.	2-propanol				isopropyl alcohol rubbing alcohol
21.			$CCl_3CH(OH)_2$		chloral hydrate sedative
22.	ethylmethanoate				rum flavor and odor

COMMON TYPES OF ORGANIC CHEMICAL REACTIONS

Combustion of Organic Compounds

$$\begin{bmatrix}\text{hydrocarbon}\\\text{alcohol}\\\text{carboxylic acid}\\\text{ester}\end{bmatrix} + O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(g)}$$

Substitution Reaction

$$\begin{bmatrix}\text{alkane}\\\text{aromatic}\end{bmatrix} + X_2 \longrightarrow \begin{bmatrix}\text{alkyl halide}\\\text{aromatic halide}\end{bmatrix} + HX$$

Addition Reactions

$$\text{alkene} + \begin{bmatrix}H_2\\HX\\X_2\\H_2O\end{bmatrix} \longrightarrow \begin{bmatrix}\text{alkane}\\\text{alkyl halide}\\\text{alkyl dihalide}\\\text{alcohol}\end{bmatrix}$$

$$\begin{matrix}\text{1 mole of}\\\text{alkyne}\end{matrix} + \text{1 mole of} \begin{bmatrix}H_2\\HX\\X_2\end{bmatrix} \longrightarrow \begin{bmatrix}\text{alkene}\\\text{alkyl halide}\\\text{alkyl dihalide}\end{bmatrix}$$

$$\begin{matrix}\text{1 mole of}\\\text{alkyne}\end{matrix} + \text{2 moles of} \begin{bmatrix}H_2\\HX\\X_2\end{bmatrix} \longrightarrow \begin{bmatrix}\text{alkane}\\\text{alkyl dihalide}\\\text{alkyl tetrahalide}\end{bmatrix}$$

Esterification

carboxylic acid + alcohol $\longrightarrow$ ester + H_2O

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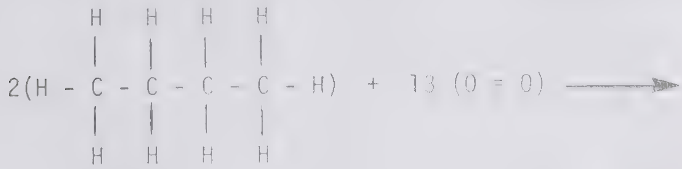
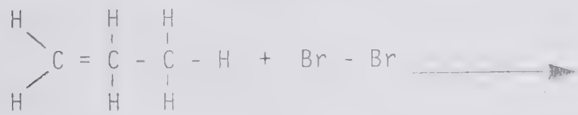
ORGANIC CHEMISTRY

Complete and balance the following chemical equations with structural formulas. [illegible] reactant and product and the reaction type.

1. 

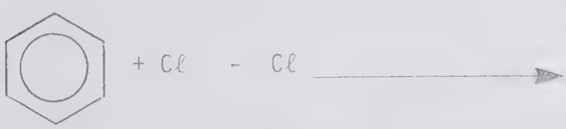
______________ + ______________ → ______________

Reaction type: ______________

2. 

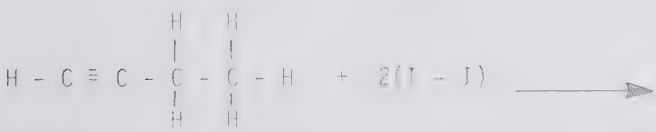
______________ + ______________ → ______________

Reaction type: ______________

3. 

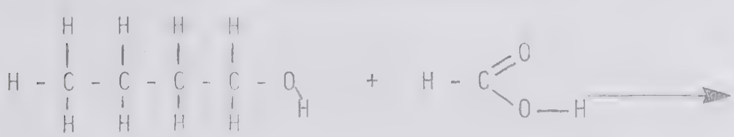
______________ + ______________ → ______________

Reaction type: ______________

4. 

______________ + ______________ → ______________

Reaction type: ______________

5. 

______________ + ______________ → ______________

Reaction type: ______________

REVIEW OF CHEMISTRY 10 & 20

ORGANIC EQUATIONS

Write the balanced structural equation for each of the following statements and name the type of reaction.

1. Methane is reacted with chlorine to make tetrachloromethane (carbon tetrachloride).

2. Ethyne is reacted with bromine to make 1,1,2,2,-tetrabromoethane (TBE), a very dense organic solvent.

3. Ethanol is made by reacting ethene with water.

4. Artificial rum flavor is made by reacting methanoic acid with ethanol.

5. Propane is burned in a camp stove.